Answer to Essential Question 15.2: An important distinction between work and the change in internal energy is that the work depends on the process involved in taking a system from one state to another, while the change in internal energy depends only on the initial and final states. Thus, if we do not know the path taken on the P-V diagram, we can not find the work – different processes have different amounts of work associated with them. On the other hand, we can find the change in internal energy, because we know the pressure and volume of the initial and final states. No matter what the process, for a diatomic ideal gas the change in internal energy in moving from state 2 to state 3 will be the +2200 J we calculated in Example 15.2B.

15-3 Constant Volume and Constant Pressure Processes

Let's consider once again two different thermodynamic processes, one in which heat is added to a system at constant volume, and the other when heat is added at constant pressure.

EXPLORATION 15.3A – A constant-volume process
A sample of monatomic ideal gas is initially at a temperature of 200 K. The gas occupies a constant volume. Heat is then added to the gas until the temperature reaches 400 K. This process is shown on the P-V diagram in Figure 15.8, where the system moves from state 1 to state 2 by the process indicated. The diagram also shows the cylinder in state 1 and again in state 2. The figure also shows the 200 K isotherm (lower) and the 400 K isotherm (higher).
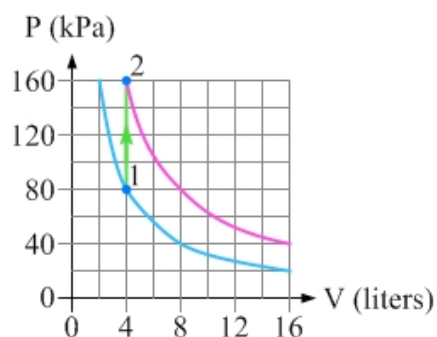

Step 1 – *Find the number of moles of gas in the cylinder.*
Applying the ideal gas law to state 1 gives:

Figure 15.8: A P-V diagram showing a constant-volume process that moves a system of monatomic ideal gas from state 1 to state 2.

$$n = \frac{PV}{RT} = \frac{(80 \text{ kPa})(4.0 \text{ L})}{(8.31 \text{ J/mol K})(200 \text{ K})} = 0.19 \text{ moles}.$$

Step 2 – *Find the work done in this process.*
The work done is the area under the curve for the process. Because there is no area under the curve in a constant-volume process the work done by the gas is zero: $W = 0$.

Step 3 – *Find the change in internal energy for this process.*
In a constant-volume process all the heat added goes into changing the internal energy of the gas. Because the gas is monatomic we have $C_V = 3R/2$. This gives:

$$\Delta E_{\text{int}} = \frac{3}{2} n R \Delta T = \frac{3}{2}(0.19 \text{ moles})(8.31 \text{ J/mol K})(400 \text{ K} - 200 \text{ K}) = +480 \text{ J}.$$

Step 4 – *Find the heat added to the gas in this process.* The First Law of Thermodynamics tells us that $Q = \Delta E_{\text{int}} + W$, but if the work done by the gas is zero we have $Q = \Delta E_{\text{int}}$. In this case we have $Q = +480$ J.

Key ideas for a constant-volume process: There is no work done by the gas: $W = 0$.

The heat added to the gas is equal to the change in internal energy: $Q = n C_V \Delta T$.

Related End-of-Chapter Exercises: 17, 18.

P (kPa)

EXPLORATION 15.3B – A constant-pressure process
      The system from the previous Exploration is now taken from state 1 to state 3 (in which, like state 2, the system is at 400 K) by the constant-pressure process indicated in Figure 15.9.

Step 1 – *Find the work done in this process.* The work done by the gas in this process is the area under the curve on the P-V diagram. Because the pressure is constant we can use Equation 15.2:

$$W = P\Delta V = (80 \text{ kPa})(8.0 \text{ L} - 4.0 \text{ L}) = (80 \text{ kPa})(+4.0 \text{ L}) = +320 \text{ J} .$$

Figure 15.9: A P-V diagram showing a constant-pressure process that moves a system of monatomic ideal gas from state 1 to state 3.

Step 2 – *Find the change in internal energy for this process.*
Because the temperature change is the same, the change in internal energy is the same as it is in the constant-volume process:

$$\Delta E_{\text{int}} = \frac{3}{2} n R \Delta T = \frac{3}{2} (0.19 \text{ moles})(8.31 \text{ J/mol K})(400 \text{ K} - 200 \text{ K}) = +480 \text{ J} .$$

Step 3 – *Find the heat added to the gas.* Applying the First Law of Thermodynamics gives:
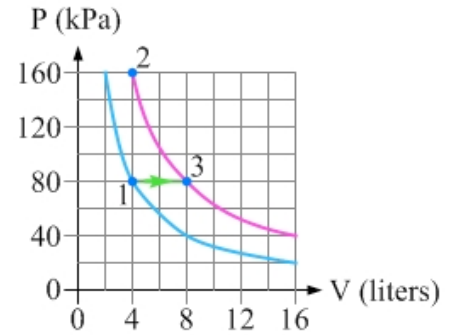$$Q = \Delta E_{\text{int}} + W = +480 \text{ J} + 320 \text{ J} = +800 \text{ J}$$

Key ideas for a constant-pressure process: The work done by the gas is given by Equation 15.2: $W = P\Delta V$.   **Related End-of-Chapter Exercises: 3, 19 – 21.**

Heat Capacity
      In Chapter 13, we used $Q = mc\Delta T$ to find the heat needed to change the temperature of a substance of mass *m* and specific heat *c*. For gases, a more convenient equation is
$Q = nC\Delta T$,   where *C* is known as the heat capacity. The value of the heat capacity depends on the process the gas follows when the heat is added. At constant volume, when the work done is zero:

$$Q = \Delta E_{\text{int}} = nC_V\Delta T , \quad \text{(Eq. 15.6: \textbf{Heat needed to change temperature at constant volume})}$$

where C_V is the heat capacity at constant volume. As mentioned in Section 15-2, we apply $\Delta E_{\text{int}} = nC_V\Delta T$ to all processes, because the change in internal energy is process independent.

Monatomic: $C_V = \frac{3}{2}R$ Diatomic: $C_V = \frac{5}{2}R$ Polyatomic: $C_V = 3R$

      In contrast, the **heat needed to change temperature at constant pressure** is given by:
$$Q = W + \Delta E_{\text{int}} = P\Delta V + nC_V\Delta T = nR\Delta T + nC_V\Delta T = n(R + C_V)\Delta T = nC_P\Delta T , \quad \text{(Eq. 15.7)}$$
      where $C_P = R + C_V$ is the heat capacity at constant pressure.

Monatomic: $C_P = \frac{5}{2}R$ Diatomic: $C_P = \frac{7}{2}R$ Polyatomic: $C_P = 4R$

Essential Question 15.3: In Exploration 15.3B, 800 J of heat added to the system of monatomic ideal gas at 200 K increases the system's temperature to 400 K via a constant-pressure process. If the same 800 J of heat is added to the system, starting in state 1, but the volume is kept constant what will the final temperature of the system be? What about the final pressure?